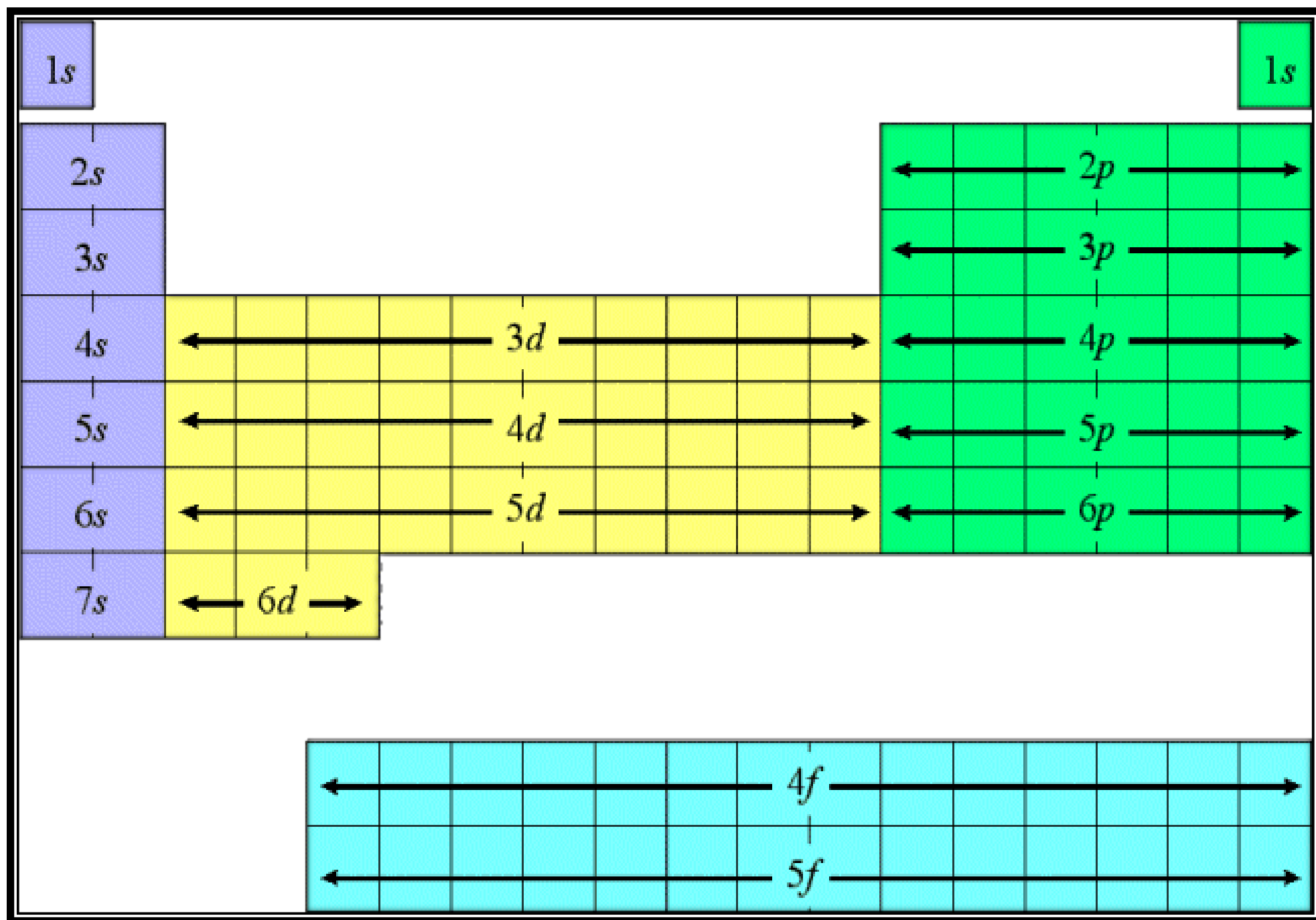


Important properties **of** **a periodic Table**

Types of Elements

Periodic Table of the Elements

H	Non-Metals																He						
Li	Be															B	C	N	O	F	Ne		
Na	Mg	Metals														Al	Si	P	S	Cl	Ar		
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr						
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe						
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn	Semi-Metals					
Fr	Ra	Ac	104	105	106	107	108	109	110	111													
			Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu							
			Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr							



1. Atomic Size

***This is the distance
between the nucleus
and the outer energy
level***

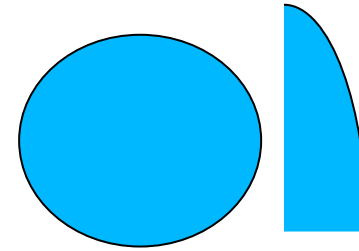
As general, the atomic size
increases on descending
a group, but decreases
across a period from left
to right.

VFI Atomic Radii

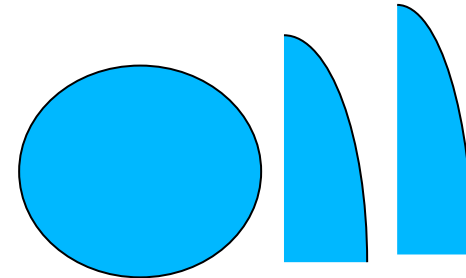


In a group

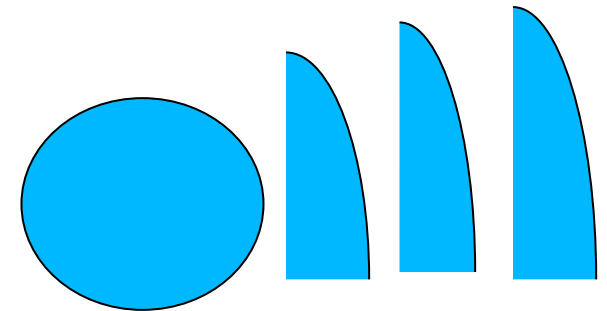
${}^1\text{H}: 1s^1$



${}^3\text{Li}: 1s^2, 2s^1$

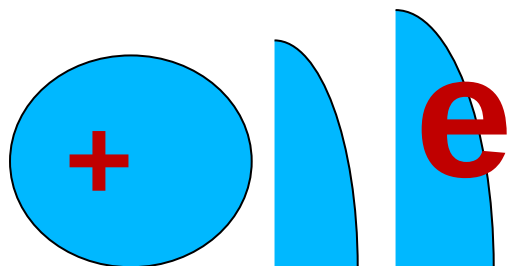


${}^{11}\text{Na}: 1s^2, 2s^2, 2p^6, 3s^1$

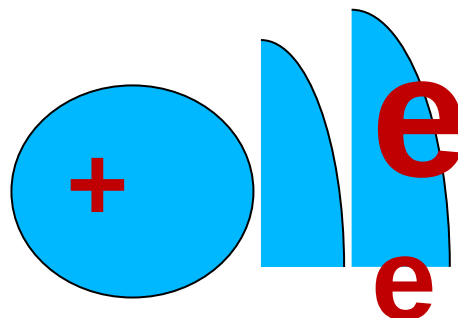


In a period

${}^3\text{Li}: 1s^2, 2s^1$



${}^4\text{Be}: 1s^2, 2s^2$



2. Ionization energy

This is the energy needed to remove an electron from a free atom of the element.

(units of ionization energy are **kJ mol^{-1}**)



As general, the ionization energies decrease on descending a group, but increase a cross a period. Although not quite regularly.

Ionization Energies

H
13.6

He
24.6

Li
5.4

Be
9.3

B
8.3

C
11.3

N
14.5

O
13.6

F
17.4

Ne
21.6

Na
5.1

Mg
7.6

Al
6.0

Si
8.2

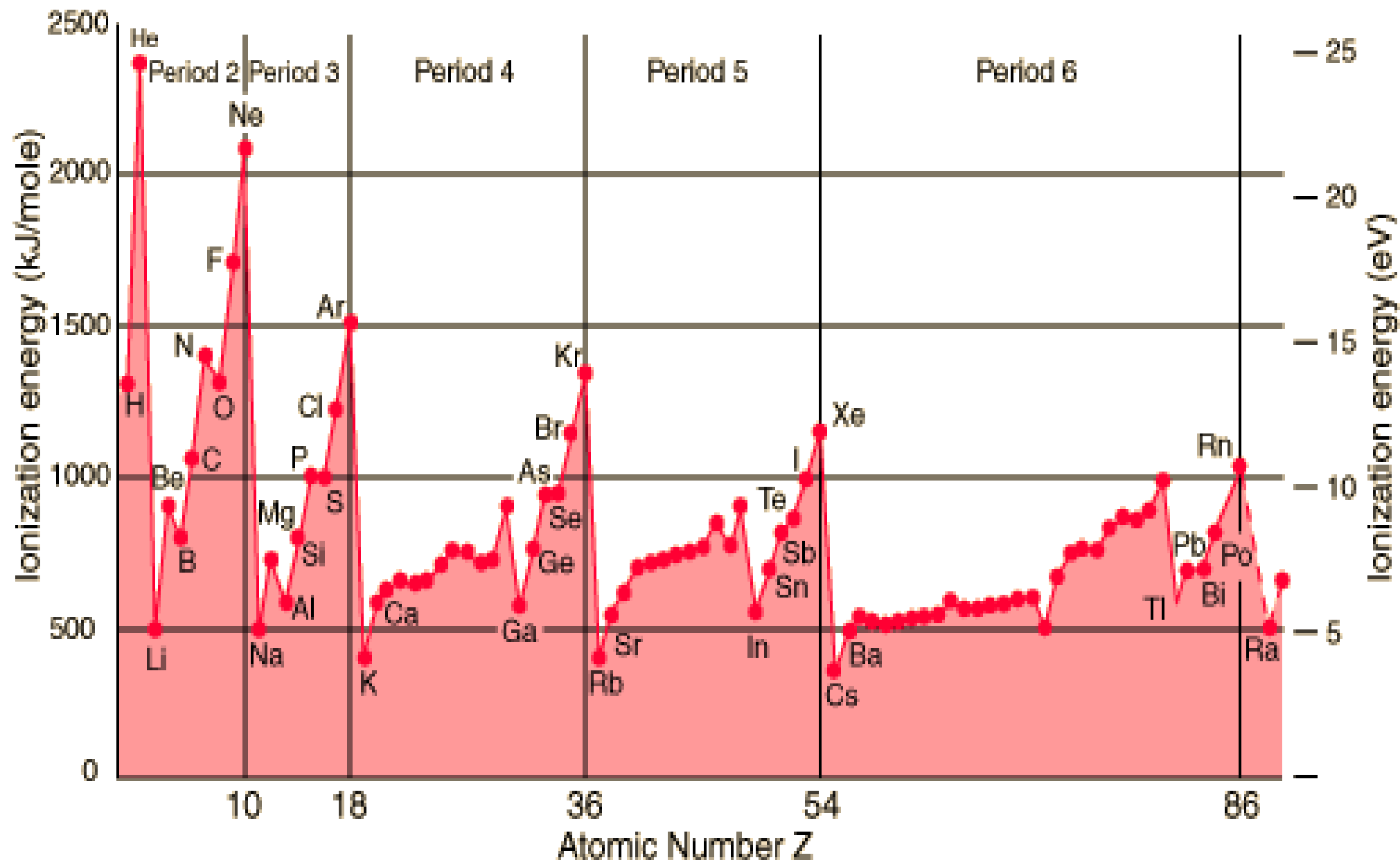
P
10.5

S
10.4

Cl
13.0

Ar
15.8

Ionization Energies



3. Electron affinity

It is the change in energy (in kJ/mole) when an electron is added to the neutral atom in the gaseous state to form a negative ion.



Electron Affinity Increases

Electron Affinity Increases

IA							VIIA	VIIIA
H 73.5							H 73.5	He *
Li 60.4	IIA			IIIA	IVA	VA	VIA	F 331.4
	Be *			B 27	C 123.4	N -7	O 142.5	Ne *
Na 53.2	Mg *			Al 45	Si 135.0	P 72.4	S 202.5	Cl 352.4
								Ar *
K 48.9	Ca *			Ga 30	Ge 120	As 78	Se 197.0	B 327.9
								Kr *
Rb 47.4	Sr *			In 29	Sn 122	Sb 102	Te 192.1	I 298.4
								Xe *
Cs 46.0	Ba *			Tl 30	Pb 110	Bi 110	Po 190	At 270
								Rn *
Fr 44.5	Ra *							

In case of stable element (As: ${}^7\text{N}$)

The element has **no** ability to **gain** or **loss** electrons

So, when we add an electron, the element must be **absorp** energy

(***Endothermic process***)



In case of unstable element (As: ${}^5\text{B}$)

The element has an ability to **gain** or **loss** electrons

So, when we add an electron, the element must be **release** energy

(***Exothermic process***)



4. Electronegativity

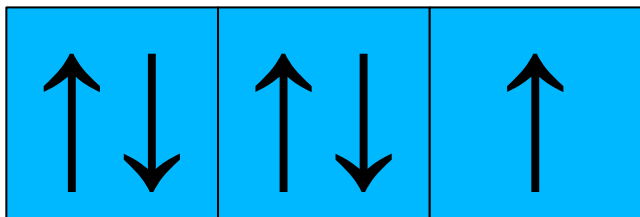
It is the ability of an atom in a compound to **attract electron** towards itself

As general, the electronegativity **decrease** on descending a **group**,
but increase a cross a **period**.

1A	2A	3B	4B	5B	6B	7B	8B	1B	2B	3A	4A	5A	6A	7A	8A		
1 H 2.1															2 He		
3 Li 1.0	4 Be 1.5										5 B 2.0	6 C 2.5	7 N 3.0	8 O 3.5	9 F 4.0	10 Ne	
11 Na 0.9	12 Mg 1.2										13 Al 1.5	14 Si 1.8	15 P 2.1	16 S 2.5	17 Cl 3.0	18 Ar	
19 K 0.85	20 Ca 1.0	21 Sc 1.3	22 Ti 1.5	23 V 1.6	24 Cr 1.6	25 Mn 1.5	26 Fe 1.8	27 Co 1.9	28 Ni 1.9	29 Cu 1.9	30 Zn 1.6	31 Ga 1.6	32 Ge 1.8	33 As 2.0	34 Se 2.4	35 Br 2.8	36 Kr 3.0
37 Rb 0.82	38 Sr 1.0	39 Y 1.2	40 Zr 1.4	41 Nb 1.6	42 Mo 1.8	43 Tc 1.9	44 Ru 2.2	45 Rh 2.2	46 Pd 2.2	47 Ag 1.9	48 Cd 1.7	49 In 1.7	50 Sn 1.8	51 Sb 1.9	52 Te 2.1	53 I 2.5	54 Xe 2.6
55 Cs 0.79	56 Ba 0.9	57 La 1.1	72 Hf 1.3	73 Ta 1.5	74 W 1.7	75 Re 1.9	76 Os 2.2	77 Ir	78 Pt 2.2	79 Au 2.4	80 Hg 1.9	81 Tl 1.8	82 Pb 1.9	83 Bi 1.9	84 Po 2.0	85 At 2.2	86 Rn 2.4
87 Fr 0.7	88 Ra 0.9	89 Ac 1.1	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Uuu	112 Uub	113 Uut	114 Uuq	115 Uup	116 Uuh		

What is the Most Electronegative Element?

Answer: Fluorine is the most electronegative element. Fluorine has an electronegativity of 4



4. Electropositivity

The opposite of electronegativity is **electropositivity**: a measure of an element's ability **to donate** **electrons**.

Magnetic Measurements

to calculate number of unpaired electrons

Paramagnetic substances:

- Contains unpaired electrons
- Drawn to magnetic field
- Magnetic moment increase as no. of unpaired electrons increase

Diamagnetic substances:

- All electrons are paired
- Not attracted to magnetic field

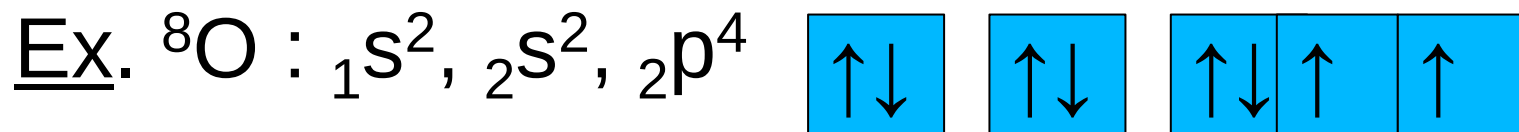
Orbital Filling & Hund's Rule

1. Electron configuration:

is the arrangement of the electrons in an atom

2. Hund's Rule:

“Electrons are distributed among the orbitals of a subshell singly then paired.”

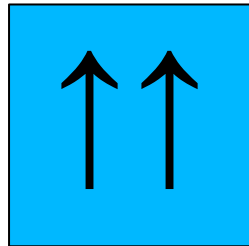


Outermost shell is called valence shell

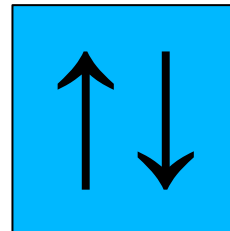
Electrons in valence shell are called valence electrons

3. Pauli exclusion principle:

The max. Number of electrons in any orbit equal two electrons & these two electrons spin in opposite direction.

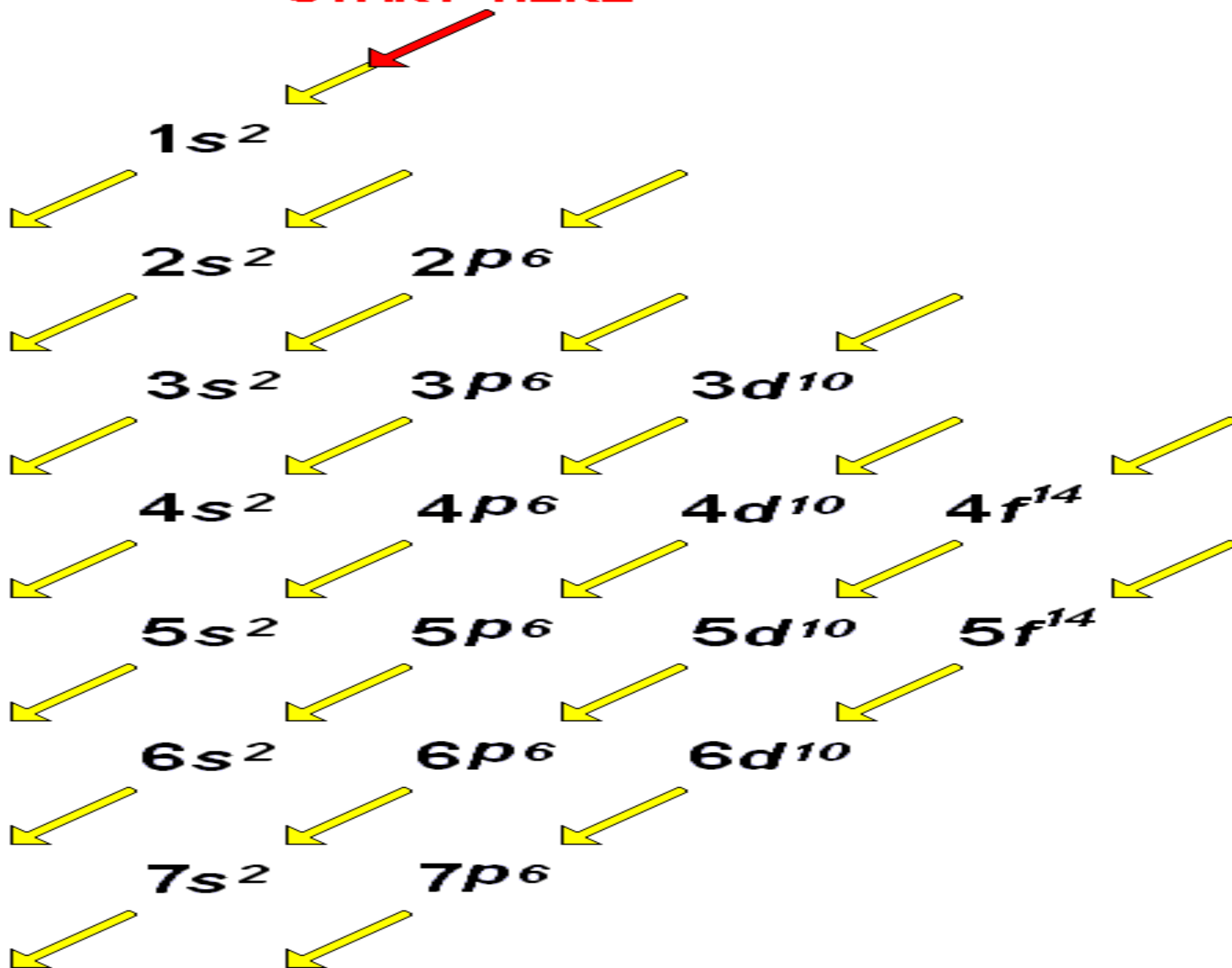


X



✓

**FOLLOW THE YELLOW BRICK ROAD --
START HERE**



❑ Method for extraction and separation:

1. Mechanical Separation:

- This process used for elements that exist in native form
- **Ex.** Diamond, Gold

2. Thermal decomposition:

➤ Mond process:



3. Substitution reaction:



4. High temperature chemical reduction method:

i. By using carbon as reducing agent



ii. By using aluminium as a highly electropositive



ii. By using electropositive metal



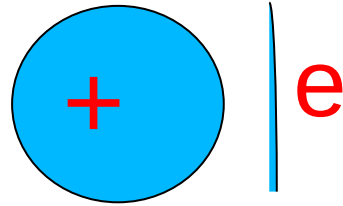
5. Electrolytic Reduction:

This method gives pure element, but it is very expensive

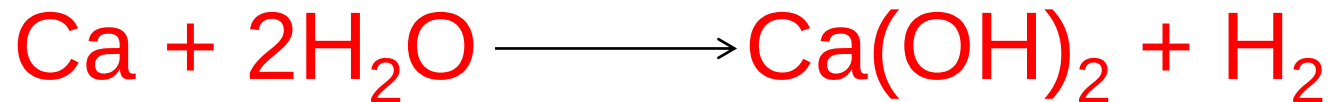
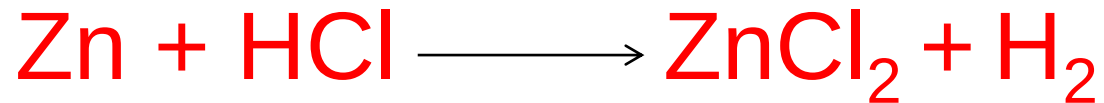
Ex: Cu , Zn

1. Hydrogen Atom

The simplest element in periodic table $1S^1$.



- Preparation of hydrogen in laboratory



- Preparation of hydrogen in Industry



❑ Isotopes:

Atoms of the same element which have the same atomic number but differ in the mass number

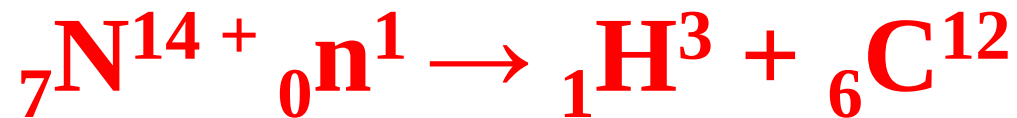
There are three types of hydrogen isotopes which are:

1. ${}_1\text{H}^1$ (hydrogen)
2. ${}_1\text{D}^2$ (deuterium)
3. ${}_1\text{T}^3$ (tritium)

1. ${}_1^3\text{T}$ (tritium) is a radioactive and emits a beta particles



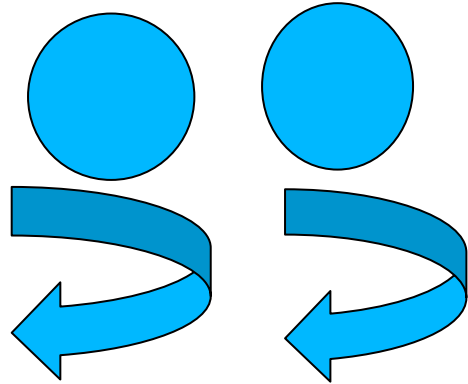
2. ${}_1^3\text{T}$ (tritium) can be prepared by:



❑ Ortho and para hydrogen

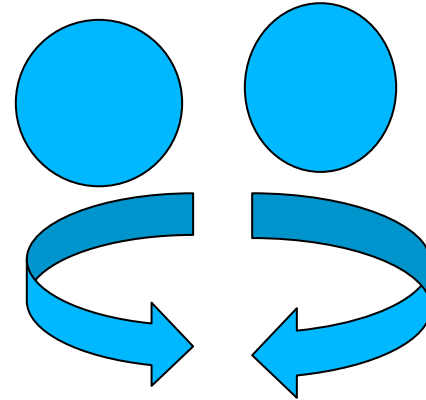
Ortho

(in the same direction)



Para

(in opposite direction)



The ortho and para hydrogen molecules differ in physical properties (b.p., heat content, thermal conductivity) **but** have the same chemical properties

❑ The uses of hydrogen

1. Preparation of ammonia



2. Preparation of water



3. Preparation of organic compounds



Group (I)

Element	Electronic configuration
Li	$[\text{He}]_2 2\text{S}^1$
Na	$[\text{Ne}]_{10} 3\text{S}^1$
K	$[\text{Ar}]_{18} 4\text{S}^1$
Rb	$[\text{Kr}]_{36} 5\text{S}^1$
Cs	$[\text{Xe}]_{54} 6\text{S}^1$

❑ The general properties of group (I) elements

1. The atomic size increase on descending the group
2. The ionization energy decrease on down the group
3. The electron affinity decrease down the group
4. The electronegativity decrease down the group

Qn. Write short notes on the electronic configuration and general properties of **group I** elements?

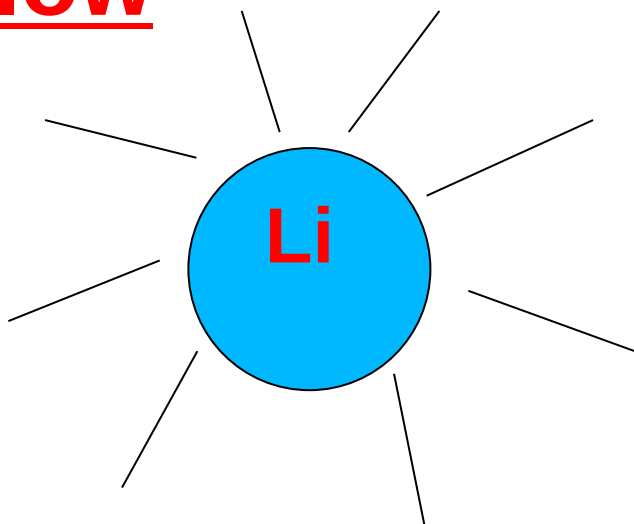
Qn. Explain the difference between the lithium and the other group I elements?

1. Li has high m.p. and B.p.
2. Li more hard
3. Li less reactive
4. Li form nitride (Li_3N) when react with air but the other group element form M_2O
5. Li is heavily hydrated

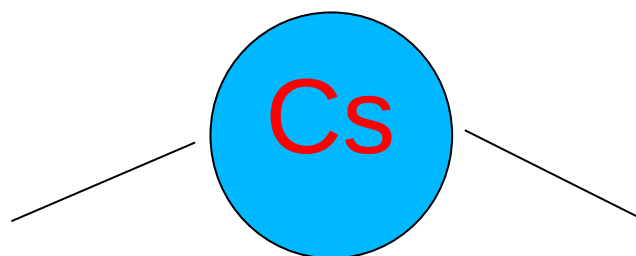
This due small size high charge

Qn. Comment on: conductivity measurements give results in order of: $\text{Cs}^+ > \text{Rb}^+ > \text{K}^+ > \text{Na}^+ > \text{Li}^+$

Slow



Fast



☐ Chemical reactions:

1. Reaction with water:



2. Reaction with O₂:



3. Reaction with H₂:



4. Reaction with N₂:



Group (II)

Element	Electronic configuration
Be	$[\text{He}]_2 2\text{S}^2$
Mg	$[\text{Ne}]_{10} 3\text{S}^2$
Ca	$[\text{Ar}]_{18} 4\text{S}^2$
Sr	$[\text{Kr}]_{36} 5\text{S}^2$
Ba	$[\text{Xe}]_{54} 6\text{S}^2$

❑ The general properties of group (II) elements

1. The atomic size increase on descending the group
2. The ionization energy decrease on down the group
3. The electron affinity decrease down the group
4. The electronegativity decrease down the group

Qn. Write short notes on the electronic configuration and general properties of **group II** elements?

❑ The general properties of group (II) elements

1. The atomic size increase on descending the group
2. The ionization energy decrease on down the group
3. The electron affinity decrease down the group
4. The electronegativity decrease down the group

Qn. Write short notes on the electronic configuration and general properties of **group II** elements?

☐ Ores

1. Be -----→ Beryl ($\text{Be}_3\text{Al}_2\text{Si}_6\text{O}_{18}$)
2. Mg -----→ Dolomite ($\text{Ca.Mg}(\text{CO}_3)_2$) & Carnalite (KMgCl_3)
3. Ca -----→ Limestone (CaCO_3) & Gypsum ($\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$)
4. Sr, Ba ---→ MCO_3 , MSO_4

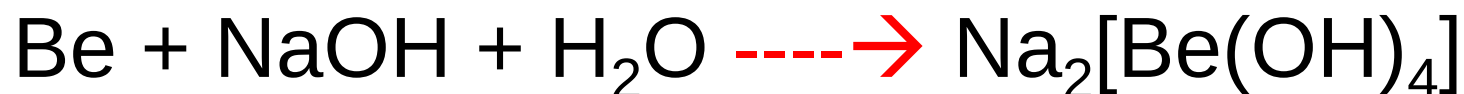
Explain the difference between the Be and the other group II elements?

1. Be has high m.p.
2. Be high electronegativity
3. Be very hard
4. Be form chelate compounds with oxygenated ligand

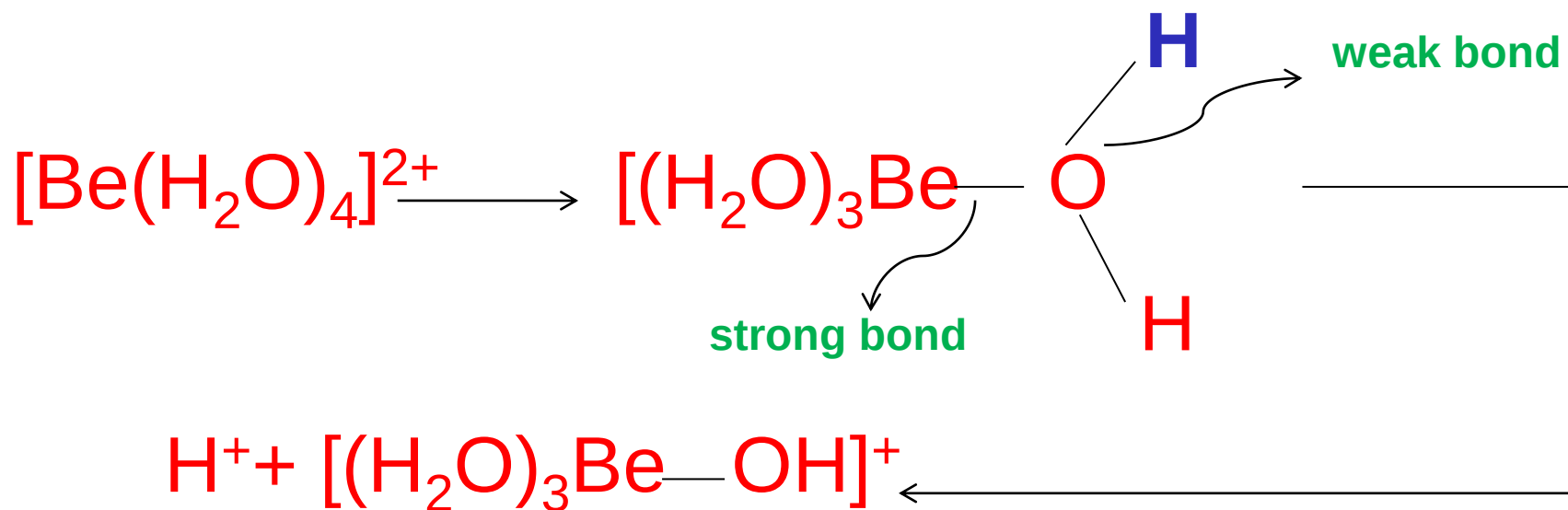
This due small size high charge

Qn. Comment on: Be metal is amphoteric?

Because it react with both acids and bases

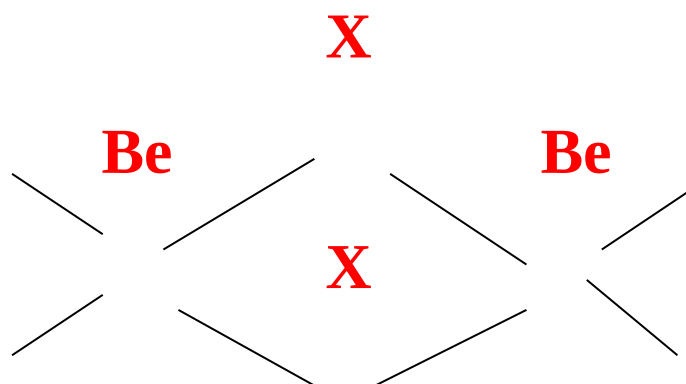


Qn. Comment on: Be salts are acidic in pure water?

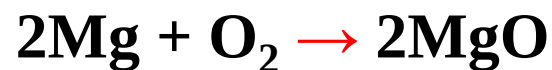
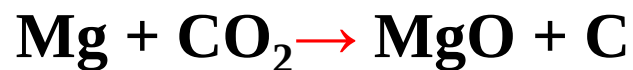


☐ Chemical reactions:

1. Reaction of Be:



2. Reaction of Mg:



(Mg used in warfare & flash bulbs, due to vigorously reaction with oxygen)

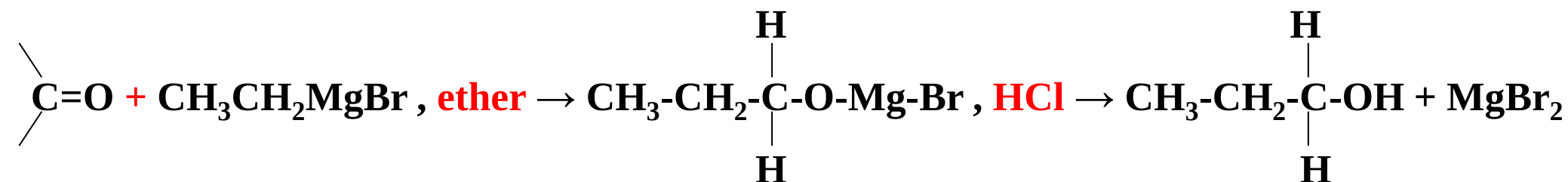


white powder

2. Preparation of Grignard reagent :



Used to synthesize alcohol from carbonyl compounds



Propanol

☐ Uses of Mg

- Alloys
- Reducing agent in extraction of U
- Containers for nuclear fuels

Group (III)

Element	Electronic configuration
B	$[\text{He}]_2 2\text{S}^2, 2\text{P}^1$
Al	$[\text{Ne}]_{10} 3\text{S}^2, 3\text{P}^1$
Ga	$[\text{Ar}]_{18} 3\text{d}^{10}, 4\text{S}^2, 4\text{P}^1$
In	$[\text{Kr}]_{36} 4\text{d}^{10}, 5\text{S}^2, 5\text{P}^1$
Tl	$[\text{Xe}]_{54} 4\text{f}^{14}, 5\text{d}^{10}, 6\text{S}^2, 6\text{P}^1$

❑ The general properties of group (III) elements

1. The atomic size increase on descending the group
2. The ionization energy decrease on down the group
3. The electron affinity decrease down the group
4. The electronegativity decrease down the group
5. Boron is semimetal and the other elements are metal
6. Metallic character increase down the group
7. The most common oxidation state is +1 & +3

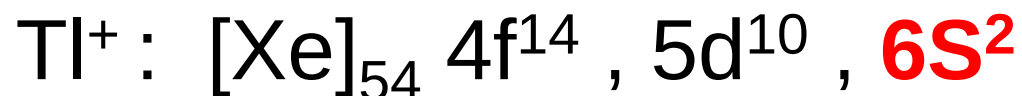
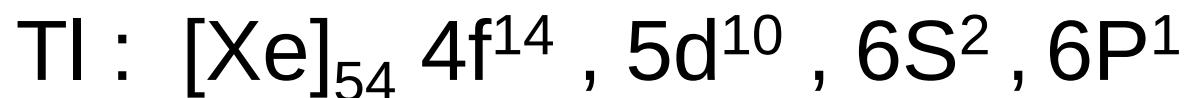
Qn. Write short notes on the electronic configuration and general properties of group III elements?

❑ Inert pair Effect:

It is the ability of the electrons in the outermost **S** orbital to remain unshared in chemical reactions, because they need high ionization energy to become unpaired.

Qn. Comment on: Thallous compound are stable

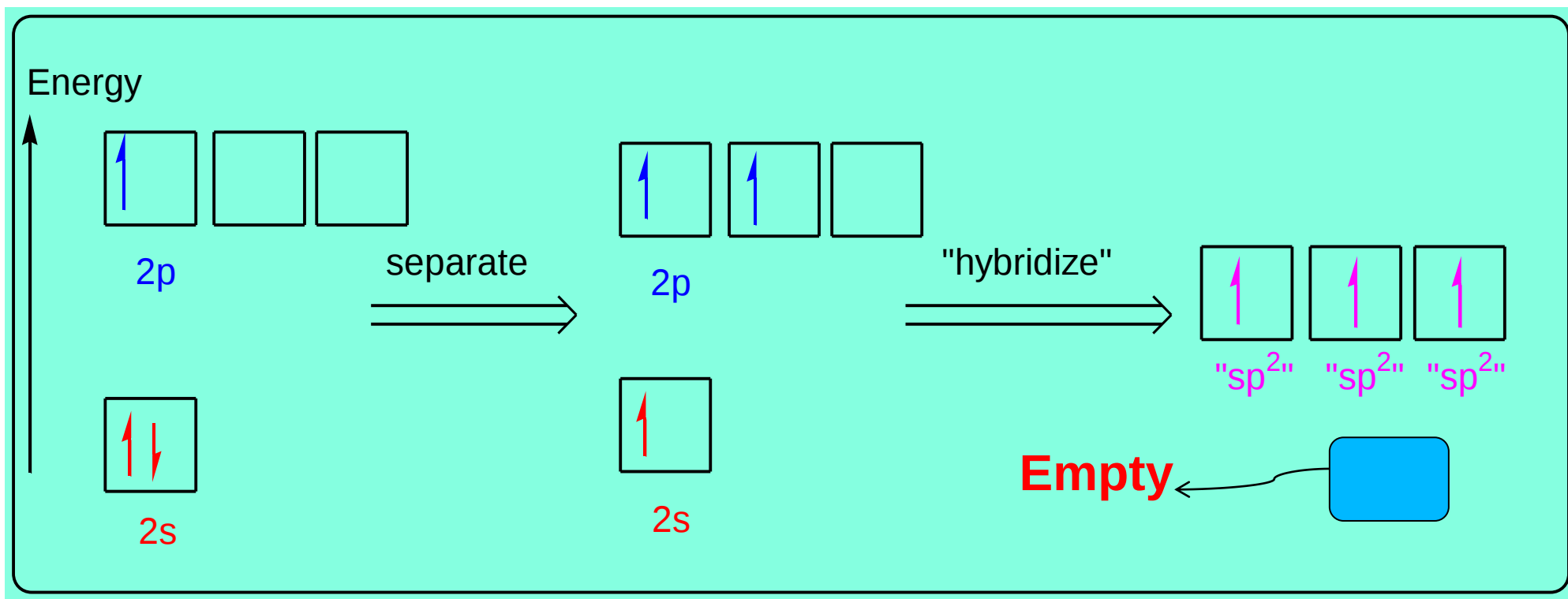
This due to the inert pair effect “It is the ability of the electrons in the outermost **S** orbital to remain unshared in chemical reactions, because they need high ionization energy to become unpaired”.



Qn. Comment on: Boron trihalid are lewis acid

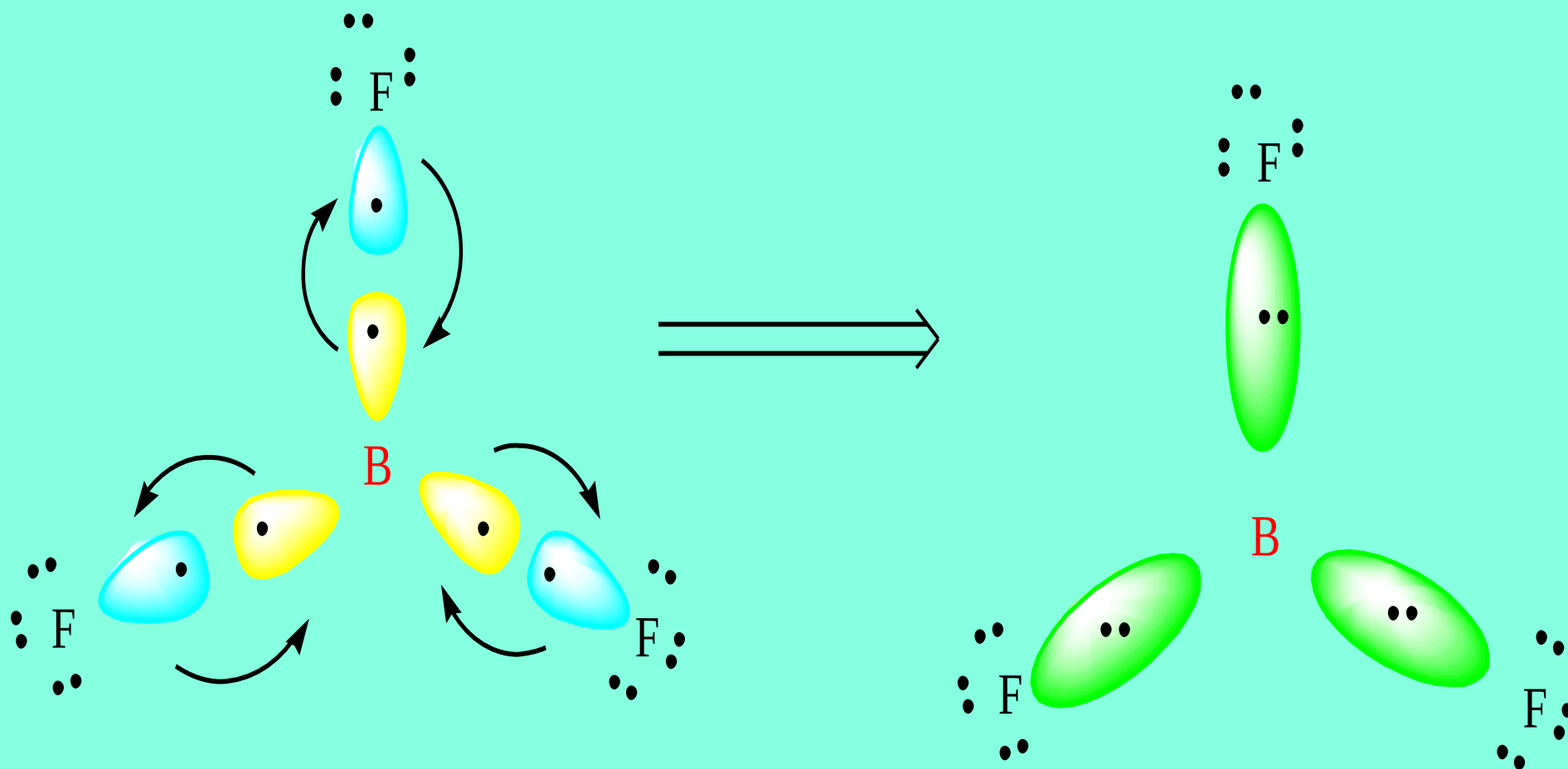
- Example: BF₃

Atomic ⁵B : 1s² 2s² 2p¹



Atomic ⁹F : 1s² 2s² 2p⁵

sp^2 hybridized, TRIGONAL PLANAR,
 120° bond angles



Qn. Comment on: Boric acid is weak monobasic - Lewis acid

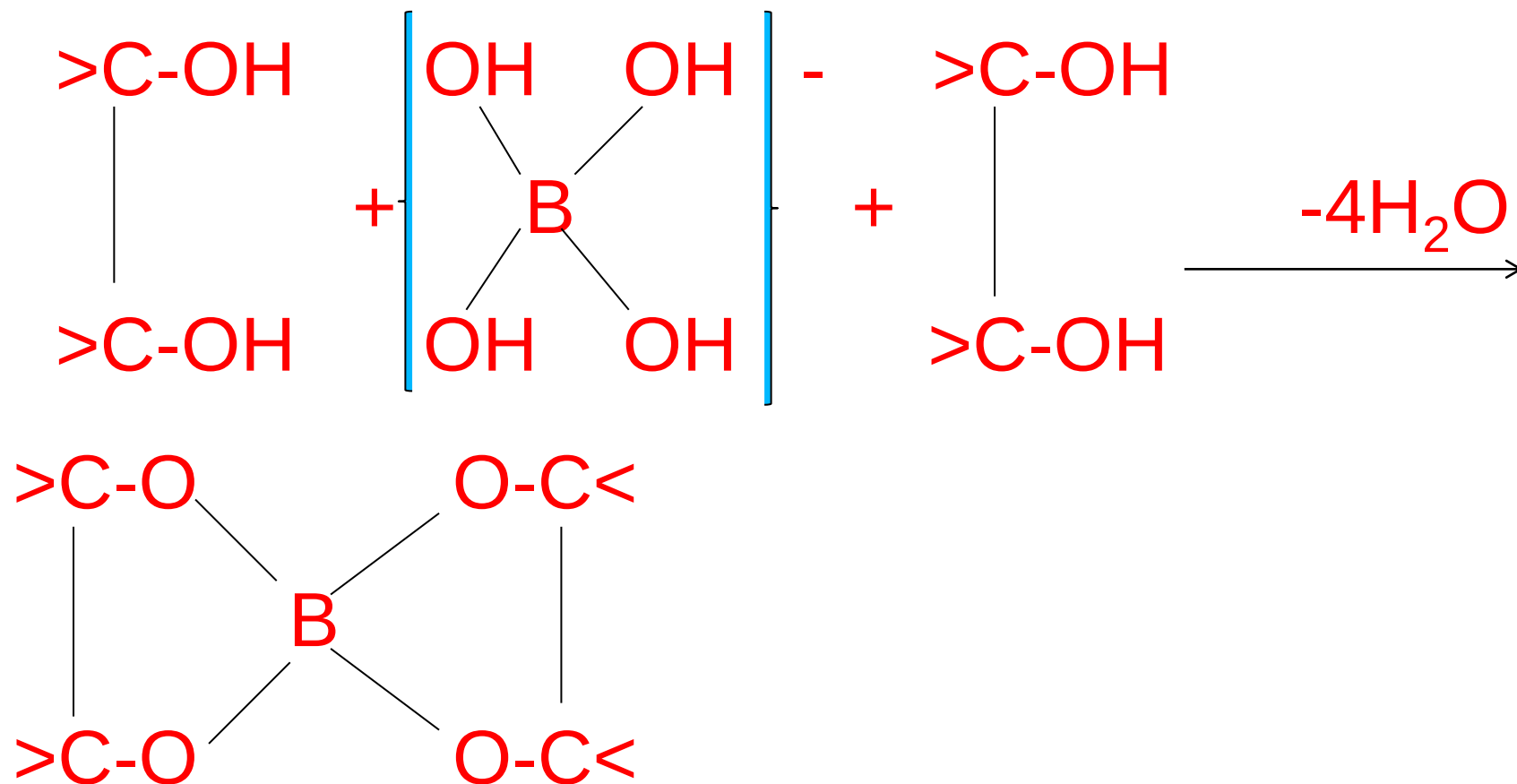
- Firstly, a monobasic acid is an acid that can give a proton (H^+ ion) in an acid-base reaction.
- A Lewis acid is a molecular unit that accepts an electron pair from another unit.
- The word "acid" in Boric acid (H_3BO_3) gives a wrong idea that Boric acid releases protons.
- Boric acid is an acid that **accept electrons** and **does not release protons**.



We see that Boric acid forms a borate ion, by accepting the electron-pair of OH^- , So, it being a Lewis acid.

The second point is that a H^+ ion is generated, which would be used up in an acid base reaction. So, it is monobasic too.

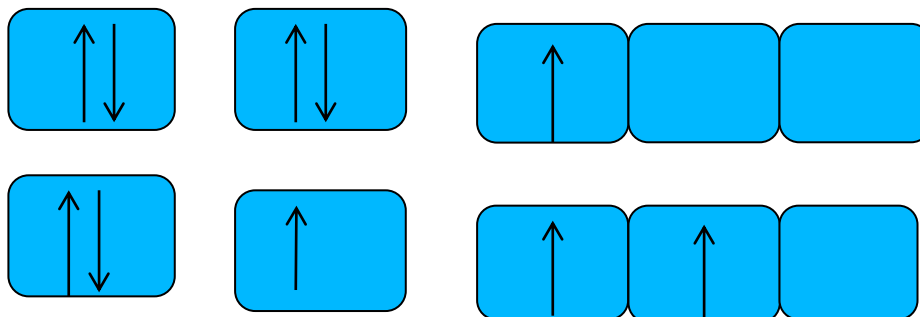
Qn. Comment on: The addition of glycerol makes boric a strong monobasic acid



Qn. Describe the structure and nature of bonding of diborane (B_2H_6)

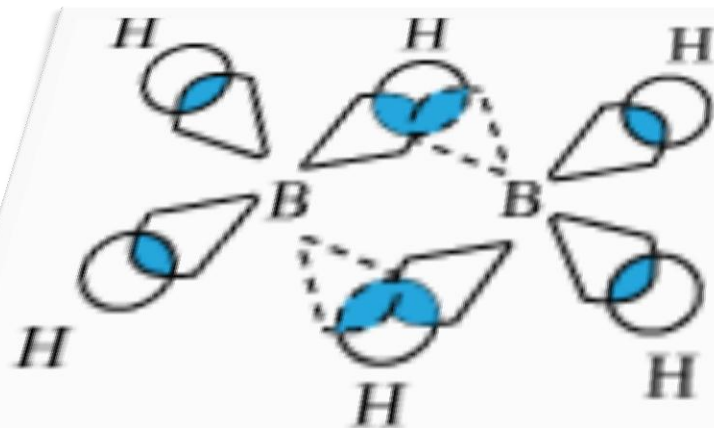


$^5B : 1S^2 2S^2 2P^1$



SP³-hyberdization
Hydrogen bridge
Banana shape

$^1H : 1S^1$

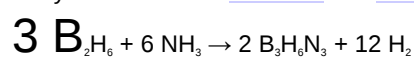


□ Boric oxides (B_2O_3)

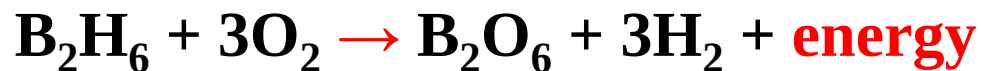


■ Properties:

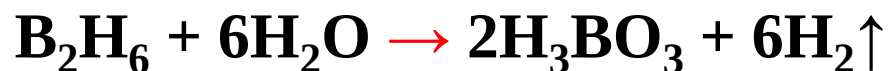
- Used to make hard resistant glass ware as pyrex
- Used as fire resistant additives for paints
- Non-metallic oxides
- Amphoteric oxides



❑ Reaction of Diborane



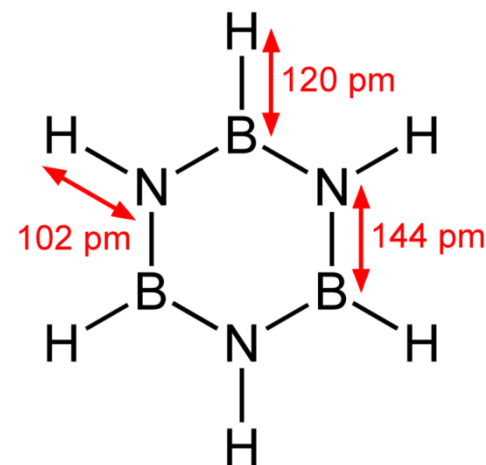
Diborane used as fuel



Used as a reducing agent



1 2



❑ Chemical properties of Al

▪ Used in Thermite reaction

Where **Al** has an ability to reduce many metallic oxides to corresponding metal

e.g. [**Cr₂O₃, Fe₂O₃, Mn₃O₄, MoO₃, ...**]



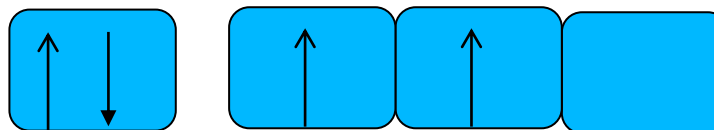
- Also, Al, is amphoteric element which it can be dissolve in both strong aqueous solution of acids and bases

Group (IV)

Element	Electronic configuration
C	$[\text{He}]_2 2\text{S}^2, 2\text{P}^2$
Si	$[\text{Ne}]_{10} 3\text{S}^2, 3\text{P}^2$
Ge	$[\text{Ar}]_{18} 3\text{d}^{10}, 4\text{S}^2, 4\text{P}^2$
Sn	$[\text{Kr}]_{36} 4\text{d}^{10}, 5\text{S}^2, 5\text{P}^2$
Pb	$[\text{Xe}]_{54} 4\text{f}^{14}, 5\text{d}^{10}, 6\text{S}^2, 6\text{P}^2$

Qn. Comment on: The limitation of carbon to coordination number 4

Because it hasn't d- orbital



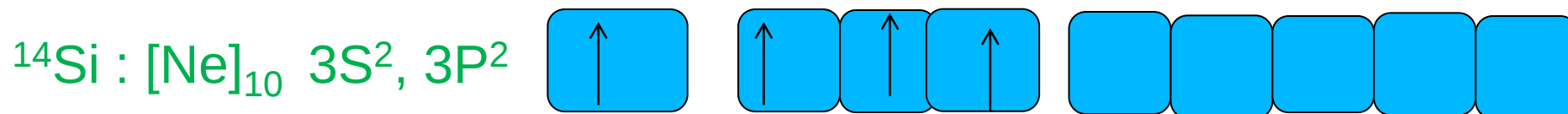
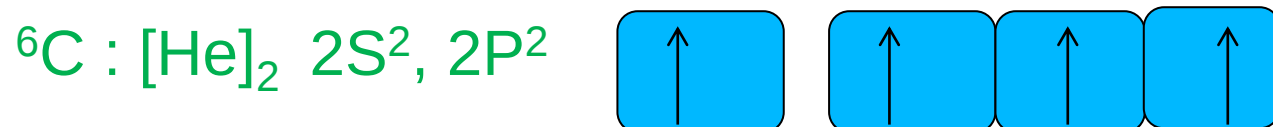
Qn. Comment on: Carbon has the ability to form multiply bond and chains

Due to the strength of the bonds, where



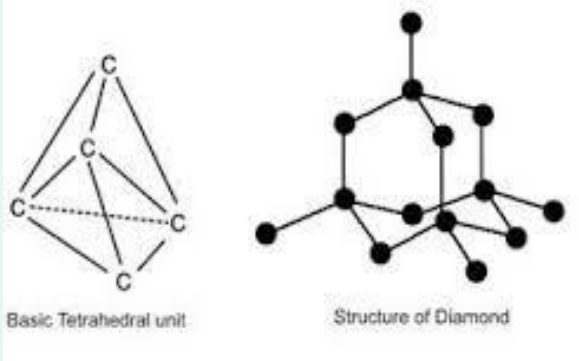
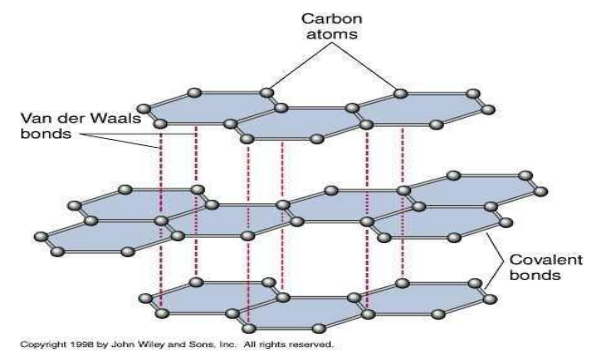
Qn. Comment on: SiF₄ can be hydrolysed while CF₄ can't be hydrolyzed with water

Because it hasn't d- orbital



☐ Allotropy phenomena:

There are a compounds which have the same atomic number but differ in the type of hybridization and the arrangement of atom in the space.

Diamond	Graphite
tetrahedral	trigonal
sp ³ -hybridized	sp ² -hybridized
109.5°	120°
hardest	soft and greasy
density is 3.5 gm/cm ³	density is 2.2 gm/cm ³
bad conductor	conduct electricity
transparent	dark gray
	



B- Carbon monoxide

Properties

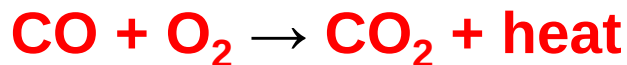
- Poisonous gas (**react with Fe in blood and prevent oxygen transfer**)
- Sparingly soluble in H₂O
- Neutral gas

Preparation

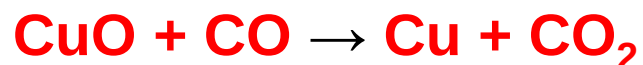
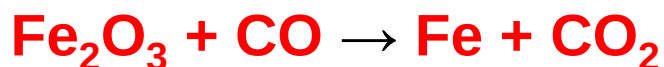


□ Uses

- Use as a fuel, because it easily oxidized

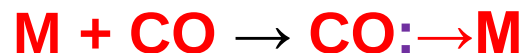


- Good reducing agent

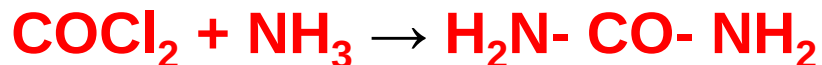


- Strong ligand and form metal complexes

CO act as a ligand and give electrons to metal



- Urea industry



3. Structure of carbon monoxide

Has linear structure (SP- Hybridization, bond angle = 180)



B- Carbon dioxide

Properties

- **CO₂ called dry ice so it used in cooling process, where it sublime at -78°C**

- **Photosynthesis process**



- **Respiration process**



- **Fermentation process**



❑ Test for CO₂

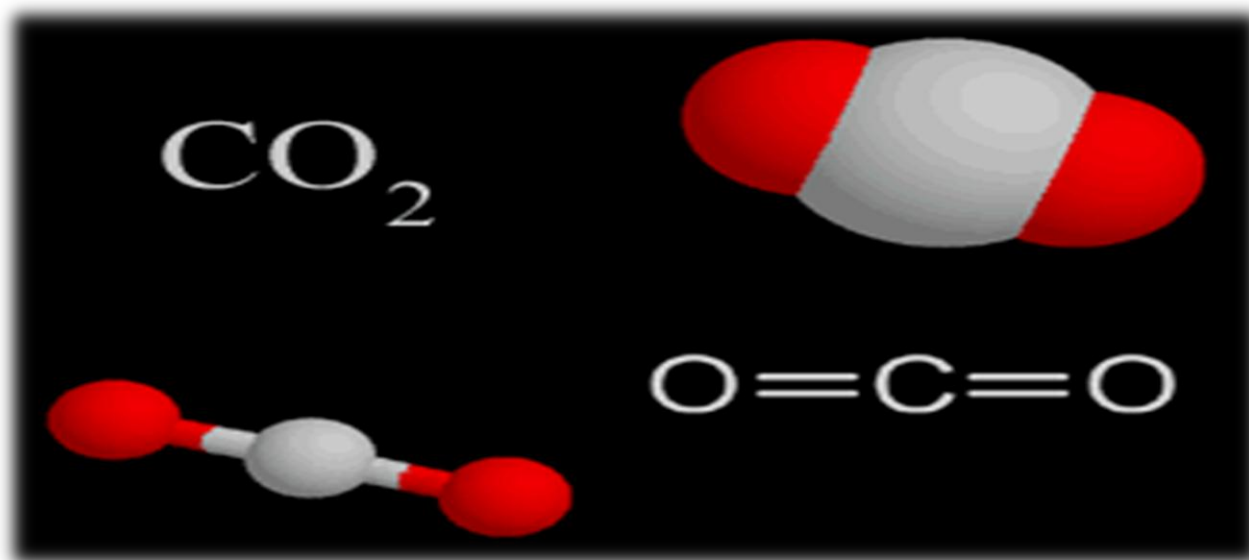
$\text{Ca(OH)}_2 + \text{CO}_2$ (**limited amount**) $\rightarrow$ CaCO_3 (**turbidity, insoluble**) $\rightarrow$ excess

$\text{CO}_2 \rightarrow \text{Ca(HCO}_3)_2$ (**turbidity disappear, soluble**)

❑ Preparation

$\text{C} + \text{O}_2$ (**excess amount**) $\rightarrow \text{CO}_2$

❑ Structure of carbon dioxide Has linear structure, sp-hybridization, bond angle 180 °



❑ Silicon

Very important element in glass, cement, transistor industries

❑ Preparation

Reduction of SiO_2 by carbon

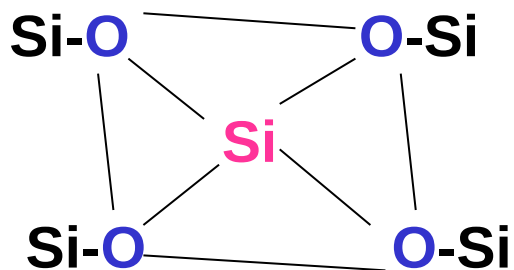


Qn. Comment on: CO_2 sublime at -78 but SiO_2 melted at 1600°C

This due to the difference in crystalline structure

$\text{CO}_2 \rightarrow$ Molecular lattice structure (**sp - hybridization**)

$\text{SiO}_2 \rightarrow$ Covalent lattice structure (**sp^3 - hybridization**)



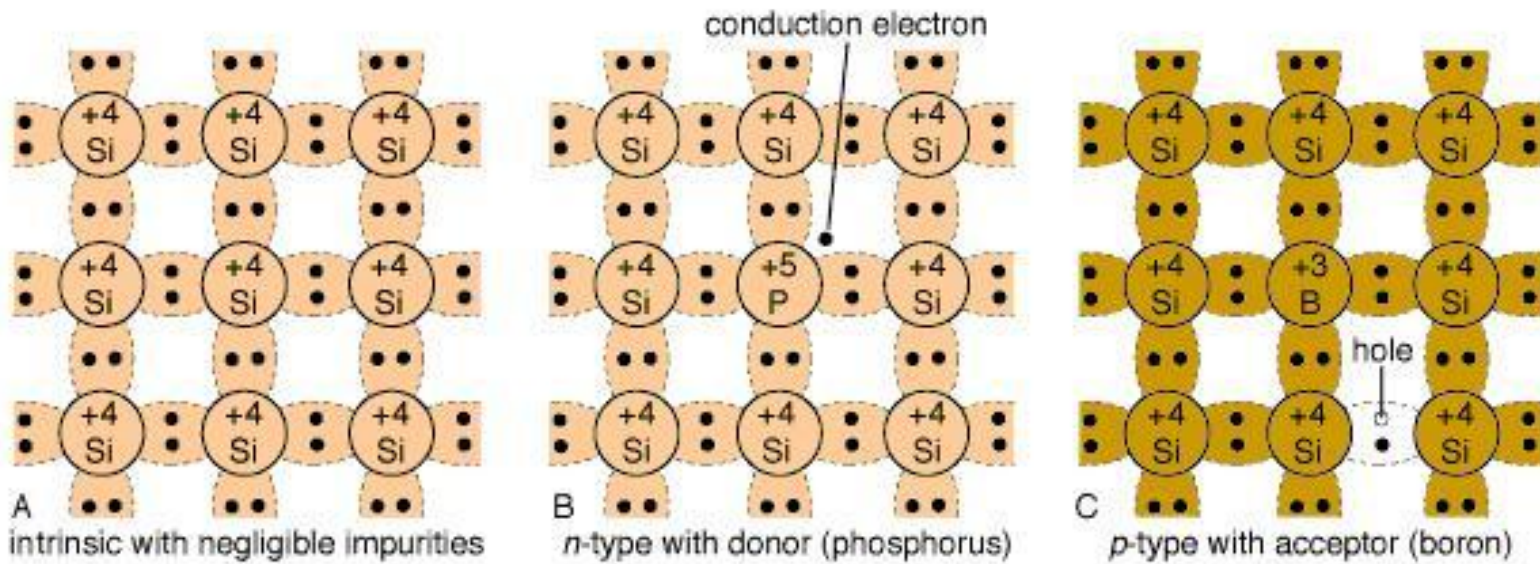
“Tetrahedral”

“N-Type Semiconductor”

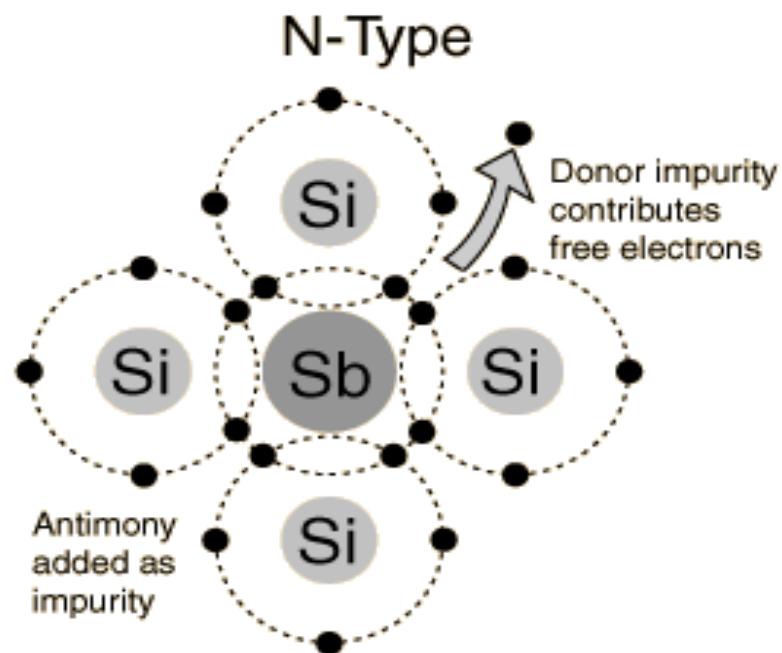
The addition of pentavalent impurities such as antimony, arsenic or phosphorous contributes free electrons, greatly increasing the conductivity of the semiconductor.

“P-Type Semiconductor”

The addition of trivalent impurities such as boron, aluminum or gallium to semiconductor creates deficiencies of valence electrons, called "holes".



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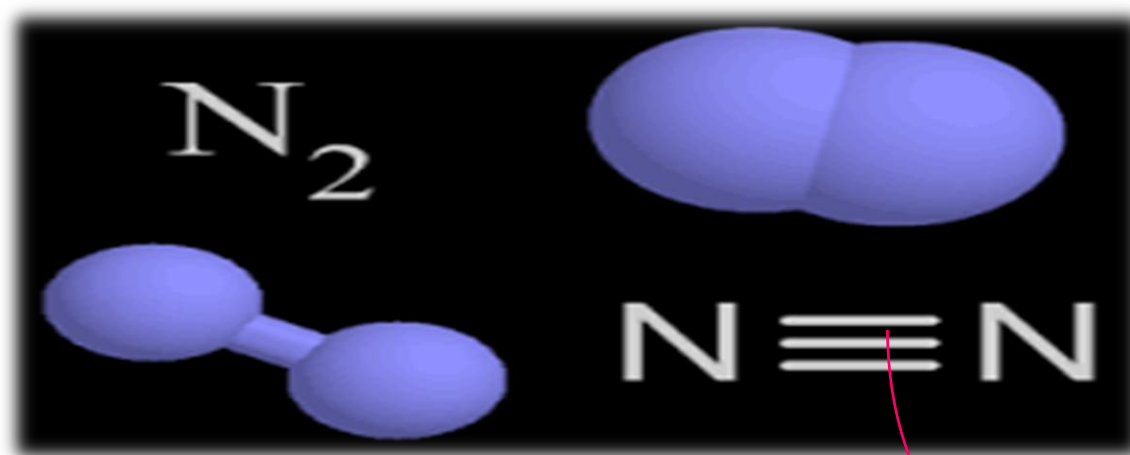
Group (V)

Element	Electronic configuration
N	$[\text{He}]_2 2\text{S}^2, 2\text{P}^3$
P	$[\text{Ne}]_{10} 3\text{S}^2, 3\text{P}^3$
As	$[\text{Ar}]_{18} 3\text{d}^{10}, 4\text{S}^2, 4\text{P}^3$
Sb	$[\text{Kr}]_{36} 4\text{d}^{10}, 5\text{S}^2, 5\text{P}^3$
Bi	$[\text{Xe}]_{54} 4\text{f}^{14}, 5\text{d}^{10}, 6\text{S}^2, 6\text{P}^3$

❑ Nitrogen compounds

1. Nitrogen molecule (N₂)

It is unreactive gas due to the bonding energy is very high of the triple bond



$\Delta H=949$

❑ Preparation of Nitrogen

1. Fractional distillation of liquified air

This method depend on the difference in the boiling point of N₂ and O₂ which B.p. of N₂ = -196 °C

B.p. of O₂ = -183 °C

2. Thermal decomposition of NH₄NO₂



Note that; NH₄NO₂ is explosive solid and the aqueous sol. is prepared by adding NH₄Cl and NaNO₂ to water and then the solution must be heated carefully to avoid explosion.

3. Passing ammonia gas over hot copper oxide



- ❑ Nitrogen fixation is a process in which nitrogen (N_2) in the atmosphere is converted into ammonium (NH_3) or nitrogen dioxide (NO_2) for example.
- ❑ [1] Atmospheric nitrogen or molecular nitrogen (N_2) is relatively inert: it does not easily react with other chemicals to form new compounds. The fixation process frees up the nitrogen atoms from their triply bonded diatomic form, $\text{N}\equiv\text{N}$, to be used in other ways.

Nitrogen fixation, natural and synthetic, is essential for all forms of life because nitrogen is required to biosynthesize basic building blocks of plants, animals and other life forms, **e.g.**, nucleotides for DNA and RNA and amino acids for proteins.

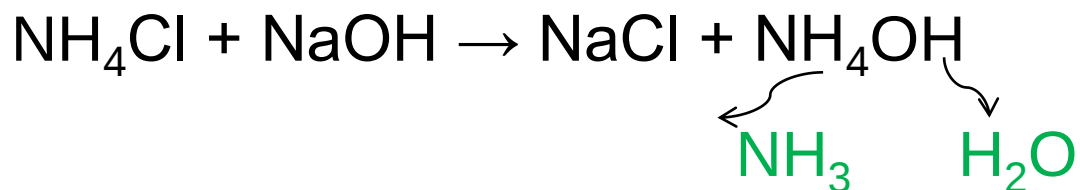
Therefore, nitrogen fixation is essential for agriculture and the manufacture of fertilizer. It is also an important process in the manufacture of explosives (e.g. gunpowder, dynamite, TNT, etc.). Nitrogen fixation occurs naturally in the air speedily.

2. Chemical fixation

I. Haber process



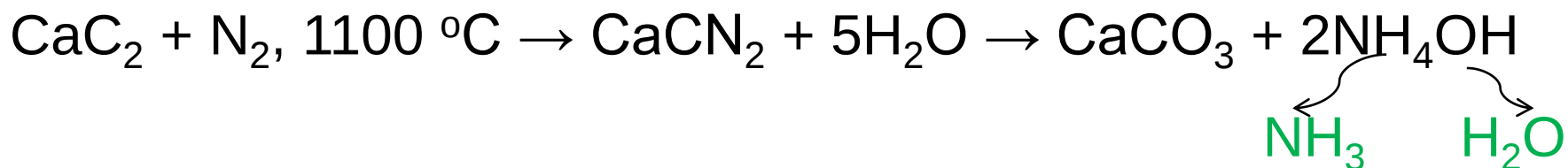
II. Substitution reaction



III. Direct reaction with lithium metal



IV. Formation of calcium cyanamide

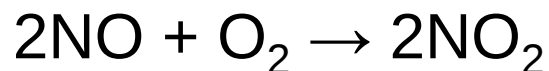


❑ Preparation of HNO₃ (Ostwald Process)

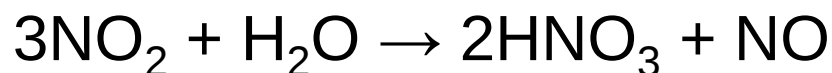
▪ Step 1: Conversion of NH₃ to NO



▪ Step 2: Conversion of NO to NO₂

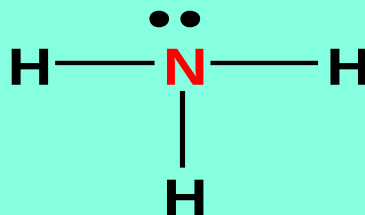
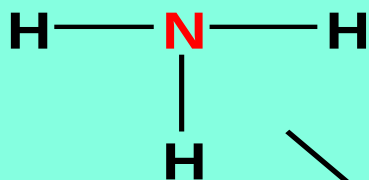


▪ Step 3: NO₂ is dissolved in water to yield HNO₃



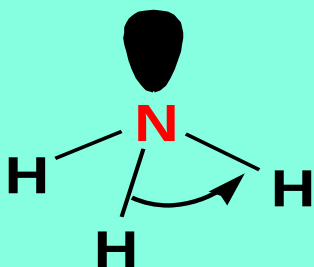
❑ Uses of HNO₃

- Strong oxidizing agent
- Production of explosive substance (**example**) trinitrotoluene (**TNT**)
- Etching and photoengraving process to produce grooves in metal surface



N	5
3H	3
8e's/2=4 prs	

Number of regions around CENTRAL ATOM: 4



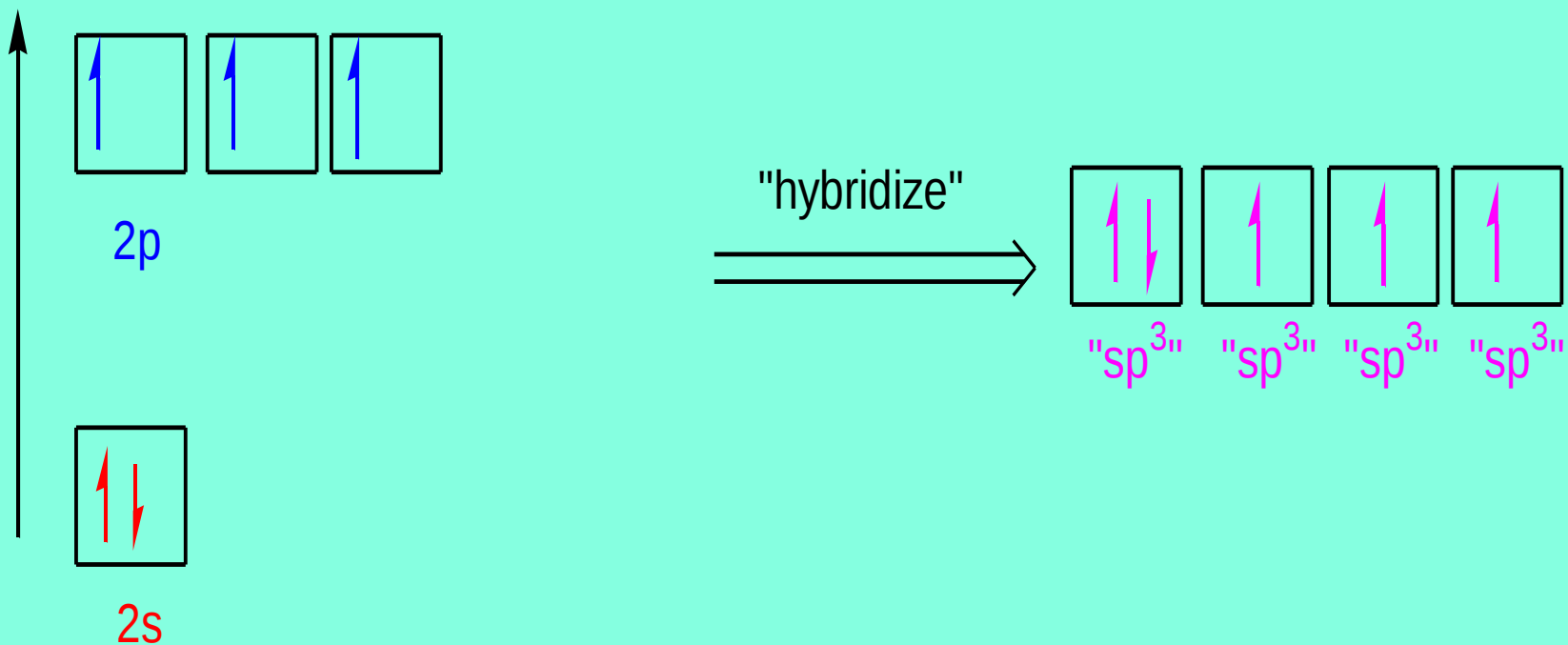
shape : **TETRAHEDRAL**
bond angles: **< 109.5°**

Hybridization of N in NH₃

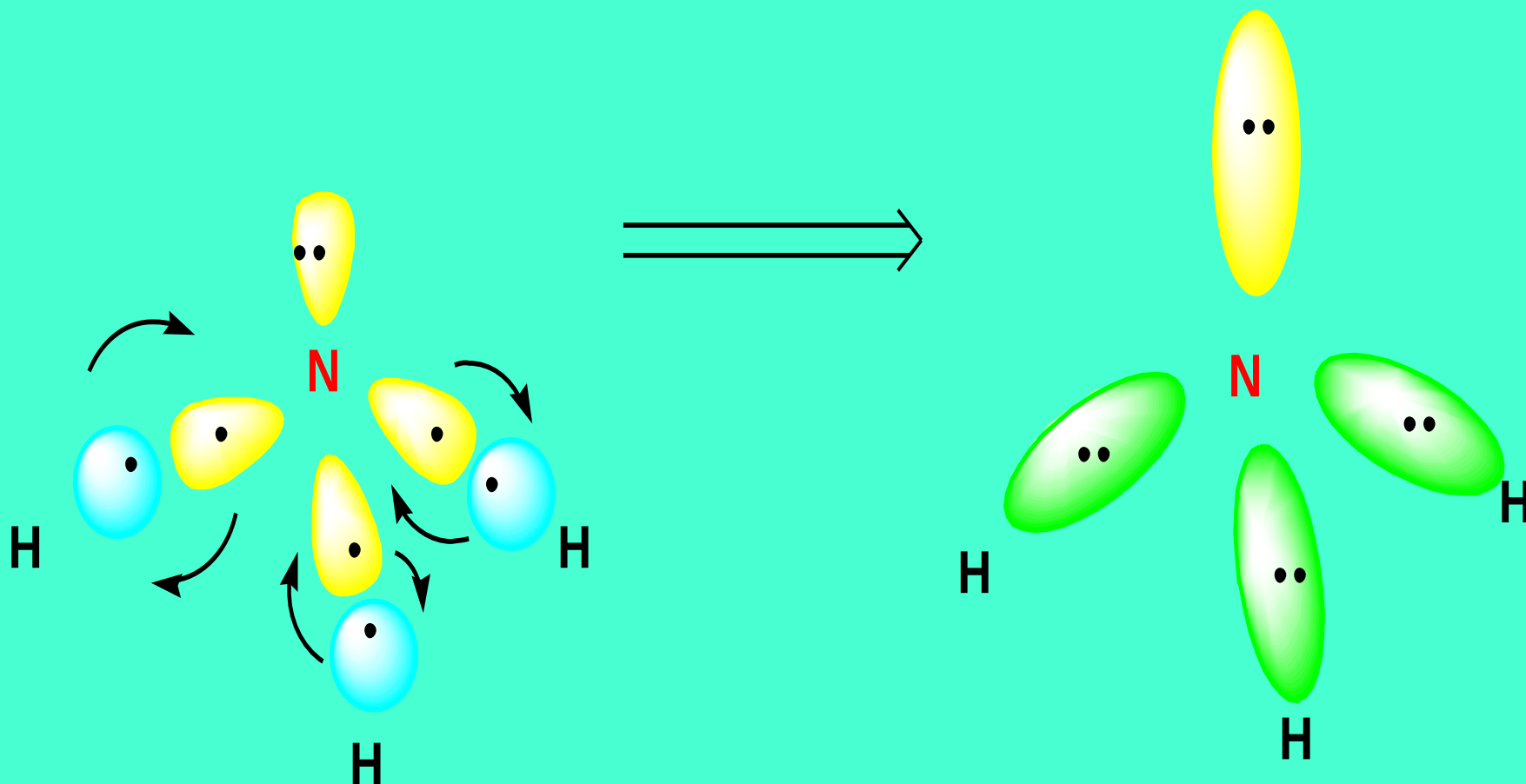
Valence e's

Atomic N: $1s^2 2s^2 2p^3$

Energy



sp^3 hybridized, TETRAHEDRAL,
~107° bond angles



Qn. Describe the structure and the nature in bonding of $\text{N}(\text{CH}_3)_3$, $\text{N}(\text{SiH}_3)_3$

► In case of $\text{N}(\text{CH}_3)_3$, nitrogen atom has sp^3 hybridization, tetrahedral

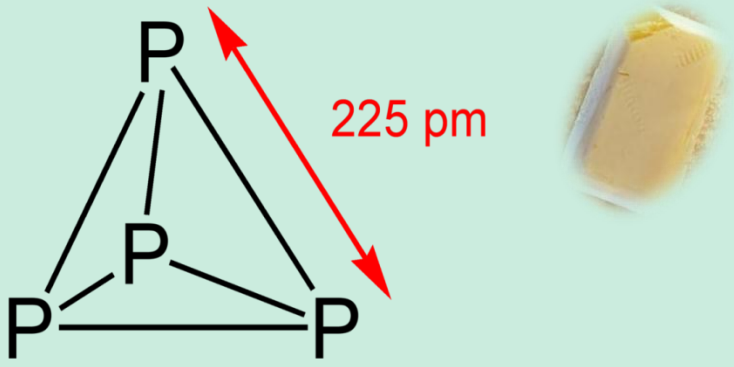
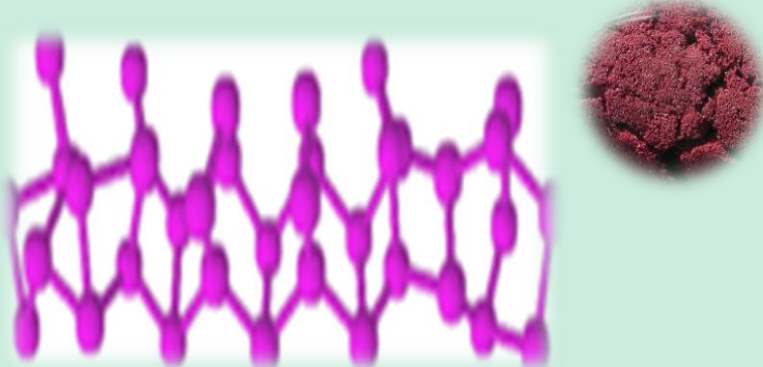
► In case of $\text{N}(\text{SiH}_3)_3$, nitrogen atom has sp^2 hybridization, triangle

Qn. Comment on: Bi in NaBiO_3 is very strong oxidizing agent

Because Bi in NaBiO_3 has oxidation state (+5) and this is unstable oxidation state So, it tend to accept electrons until it reach to stable oxidation state (+3)

☐ Allotropy phenomena:

There are a compounds which have the same atomic number but differ in the type of hyberdization and the arrangment of atom in the space.

White Phosphorus	Red Phosphorus
tetrahedral	Amorphous (random aggregates)
sp ³ -hybridized	-
Highly flammable and self-igniting upon contact with air as well as toxic causing severe liver damage	formed by heating white phosphorus to 250 °C or by exposing white phosphorus to sunlight
soluble in benzene, oils, carbon disulfide and slightly soluble in water	soft and greasy
	



Group (VI)

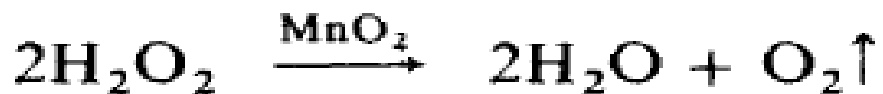
Element	Electronic configuration
O (Oxygen)	$[\text{He}]_2 2\text{S}^2, 2\text{P}^4$
S (Sulphur)	$[\text{Ne}]_{10} 3\text{S}^2, 3\text{P}^4$
Se (Selenium)	$[\text{Ar}]_{18} 3\text{d}^{10}, 4\text{S}^2, 4\text{P}^4$
Te (Tellurium)	$[\text{Kr}]_{36} 4\text{d}^{10}, 5\text{S}^2, 5\text{P}^4$
Po (polonium)	$[\text{Xe}]_{54} 4\text{f}^{14}, 5\text{d}^{10}, 6\text{S}^2, 6\text{P}^4$

Oxygen

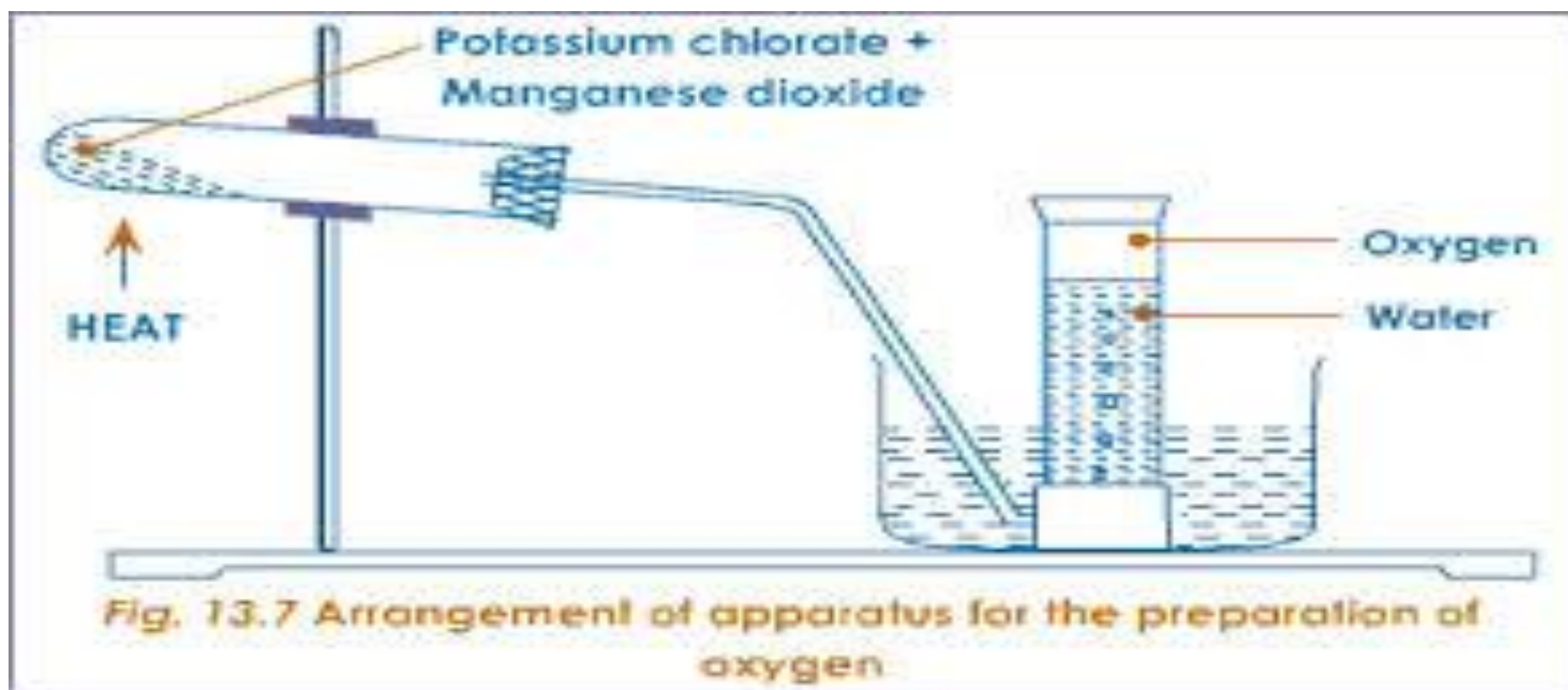
- Oxygen occurs free in the atmosphere (21 % by volume).
- Water contains 89 % by weight of oxygen
- The outer crust of the earth contains about 47 %
- Hence air, earth and sea together contain about 50% by weight of oxygen.

On the industrial scale oxygen is obtained by

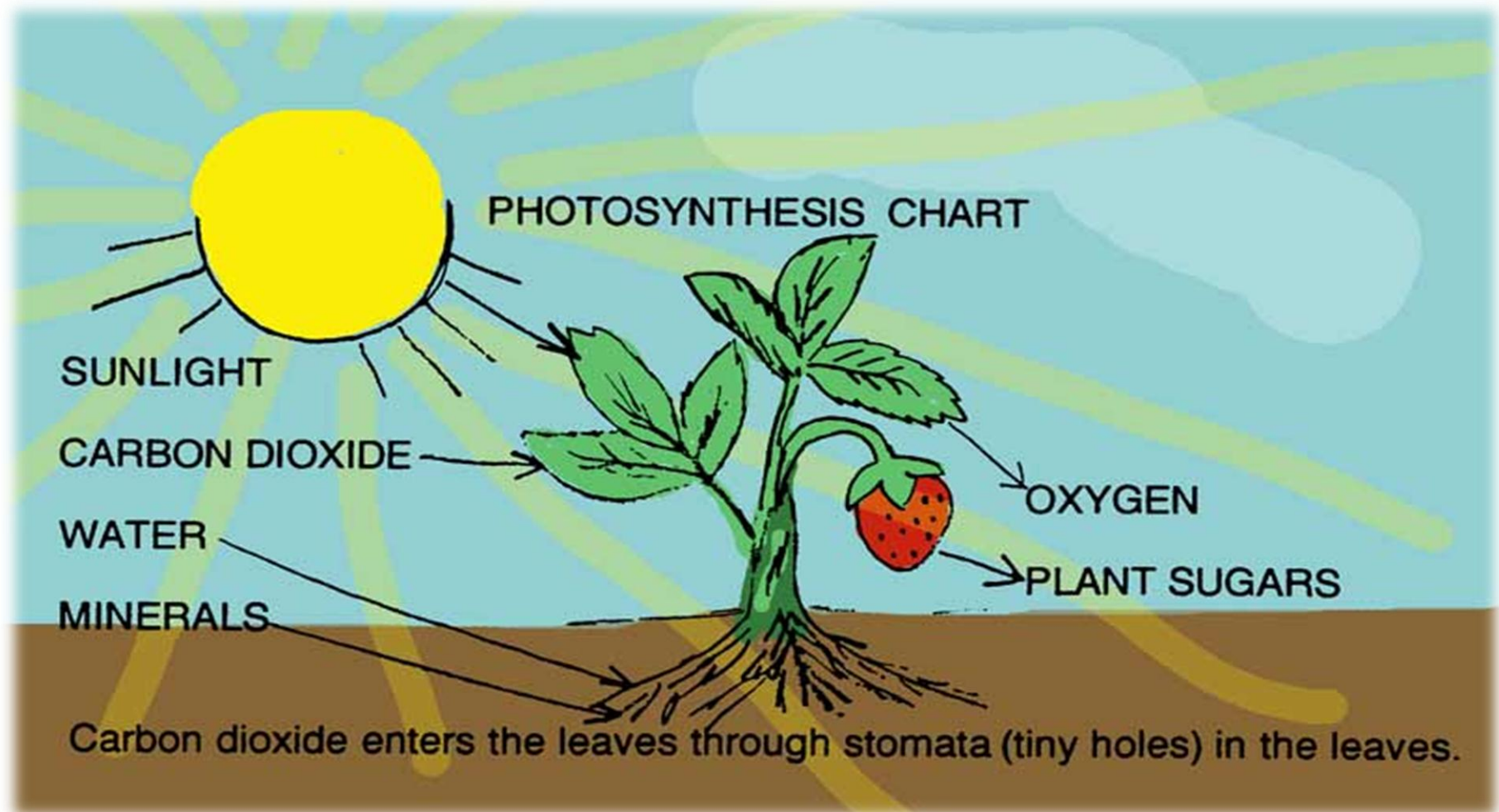
- The fractional distillation of air.
- A common laboratory method for the preparation of oxygen is by the decomposition of hydrogen peroxide. H_2O_2 , a reaction catalyzed by manganese(IV) oxide:



“Preparation of oxygen in laboratory”

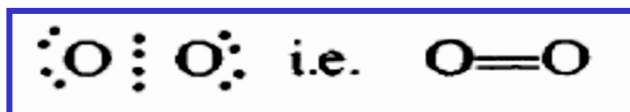


“Preparation of oxygen by Photosynthesis Process”



The allotropic forms of oxygen, oxygen, O_2 and ozone, O_3 .

Oxygen, O_2 , is a colorless gas which condenses to a pale blue liquid, b.p. 90 K, which is markedly paramagnetic indicating the presence of unpaired electrons.



Ozone, O_3 , is found in trace quantities in the upper atmosphere where it is believed to be formed by the photochemical dissociation of oxygen molecules by the intense ultra-violet light from the sun;

Small quantities of ozone are produced when oxygen and air are subjected to an electrical discharge.

Probably a small quantity of atomic oxygen is initially produced; most of this recombines quickly to give oxygen, O_2 , but a few atoms react to form ozone



❑ Uses of O₂

1. **Manufacture of steel**
2. **Hospital**
3. **Breathing under water**
4. **Formation HNO₃**
5. **Metal welding**

❑ **Other important uses include organic oxidation reactions;**

The oxidation of ethene CH₂=CH₂ to epoxyethane, CH₂(O)CH₂

❑ Uses of Sulphur

- **Is used in the manufacture of matches and fireworks**
- **As a dust insecticide**
- **Used for the manufacture of sulphuric acid**

Qn. **Explain the allotropic forms of sulfur**

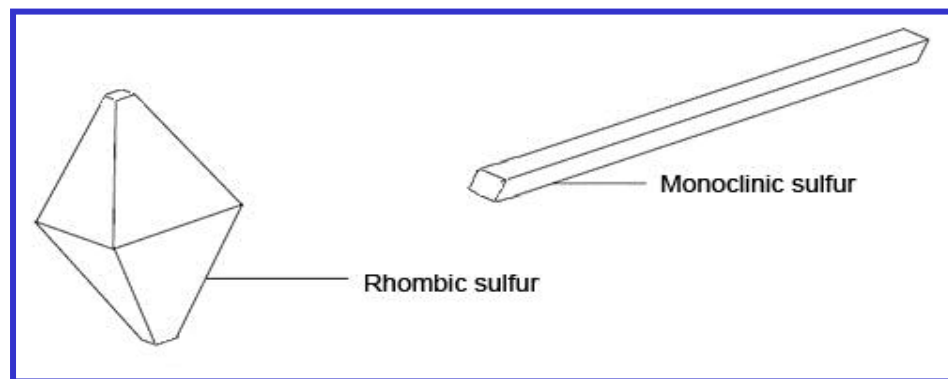
1. Rhombic sulphur $\rightarrow S_8$ (stable at R.T.)
2. Monoclinic sulphur $\rightarrow S_8$ (stable above 369 K)

The difference between them is the arrangement of atoms in the space

Rhombic Sulphur is the solid allotrope, stable at room temperature, yellow, soluble in carbon disulphide, from which it can be crystallized and has a density of 2.06 g cm^{-3} .

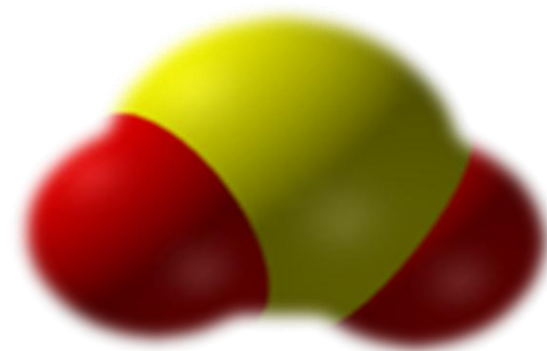
Above 369 K, rhombic Sulphur is no longer stable. slowly changing to **monoclinic**

Sulphur is long needle-like crystals (almost colourless) of monoclinic sulphur, density 1.96 g cm^{-3}

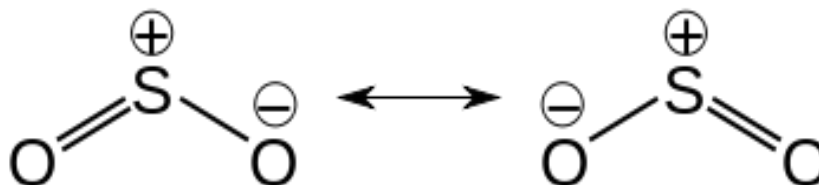


Sulfur dioxide

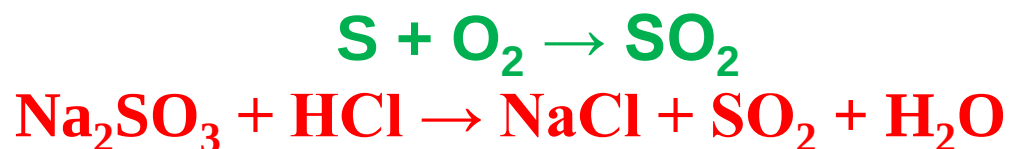
Sulfur dioxide is the chemical compound with the formula SO_2 at standard atmosphere, it is a toxic gas with a pungent, irritating, and rotten smell.



SO_2 is a bent molecule (**triangle**) with sp2 hybridization, bond angle = 95.5, which has two resonance structures

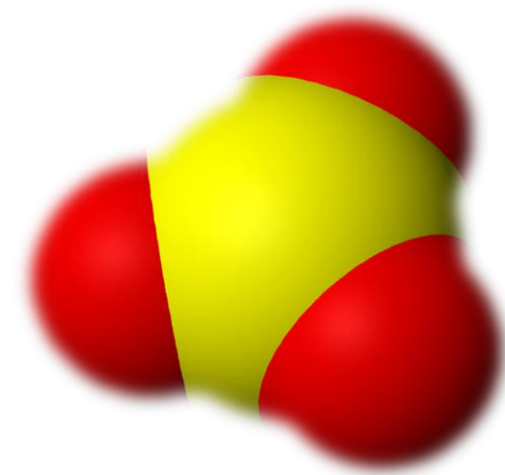


Sulfur dioxide is the product of the burning of sulfur or of burning materials that contain sulfur



Sulfur trioxide

Sulfur trioxide is the chemical compound with the formula SO_3 . In the gaseous form, this species is a significant pollutant, being the primary agent in acid rain.

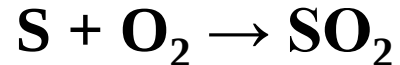


SO_3 is a Trigonal planar with sp^2 hybridization

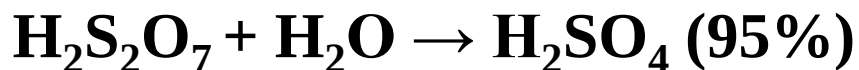
SO_3 is prepared from SO_2 , by oxidised the atmospheric oxygen at between 400 and 600 °C over a catalyst consisting of vanadium pentaoxide (V_2O_5)



❑ Preparation of H_2SO_4 by Contact's method



SO_3 absorbed into pure $\text{H}_2\text{SO}_4 \rightarrow$ Olume (**fuming sulphoric**)



❑ Uses of H_2SO_4

1. Fertilizer

2. Chemical industry

3. Synthetic fiber

4. Used as acidic medium in redox titration with KMnO_4

Group (VII)

Element	Electronic configuration
F	$[\text{He}]_2 2\text{S}^2, 2\text{P}^5$
Cl	$[\text{Ne}]_{10} 3\text{S}^2, 3\text{P}^5$
Br	$[\text{Ar}]_{18} 3\text{d}^{10}, 4\text{S}^2, 4\text{P}^5$
I	$[\text{Kr}]_{36} 4\text{d}^{10}, 5\text{S}^2, 5\text{P}^5$

❑ Halogens as ligands

- The small fluoride ion shows a great affinity to act as a ligand and form complex ions,
- For example $[\text{AlF}_6]^{3-}$, $[\text{PF}_6]^-$, $[\text{FeF}_6]^{3-}$
- But the other larger halide ions (iodide ion) show this affinity to form complexes which are usually *less stable*,
- In certain cases there is not enough space around the atom for as many iodine atoms as for other halogens, for example phosphorus forms Penta halides with fluorine, chlorine and bromine, and in the case of fluorine the ion $[\text{PF}_6]^-$, but no Penta iodide. The large size of iodine also accounts for the fact that there are few complexes with more than four iodine ligands.

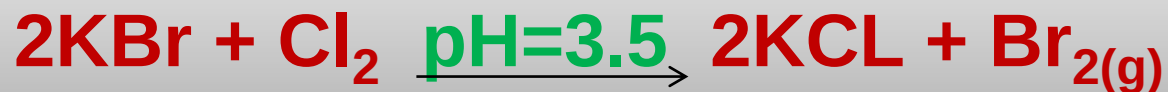
1. Preparation of fluorine (F₂)



2. Preparation of chlorine (Cl₂)



3. Preparation of bromine (Br₂)



4. Preparation of Iodine (I₂)



Qn.1 Give reason HF is a weak acid?

Due to the very strong H-F bond, where fluoride ion has high electronegative value.

Qn.2 Comment HBr & HI can not be prepared by heating NaBr or NaI with conc. H_2SO_4 ?

Because H_2SO_4 oxidize HBr to Br_2 & HI to I_2 .



Group (0)

Element	Electronic configuration
He	$1S^2$
Ne	$1S^2, 2S^2, 2P^6$
Ar	$1S^2, 2S^2, 2P^6, 3S^2, 3P^6$
Kr	$1S^2, 2S^2, 2P^6, 3S^2, 3P^6$ $, 4S^2, 3d^{10}, 4P^6$
Xe	$1S^2, 2S^2, 2P^6, 3S^2, 3P^6$ $, 4S^2, 3d^{10}, 4P^6, 5S^2, 4d^{10}, 5P^6$
Rn	$1S^2, 2S^2, 2P^6, 3S^2, 3P^6$ $, 4S^2, 3d^{10}, 4P^6, 5S^2, 4d^{10}, 5P^6$ $6S^2, 4f^{14}, 5d^{10}, 6P^6$

☐ Common properties of Noble Gases

☐ The elements classed as Noble Gases have the following properties in

common:

1. They are non-metals
2. Very un-reactive gases
3. Colorless
4. Exist as single atoms
5. Very low electro-negativities
6. High ionization energies